Student Textbook pages 526–529

Answers to Understanding Concepts Questions

1. Oxidation was originally defined as the reaction of an atom or compound with oxygen. Since then, oxidation has been redefined as a reaction in which electrons are lost.

$2Na(s) + O_2(g) \rightarrow Na_2O(s)$

ionic equation: $2Na(s) + O_2(g) \rightarrow 2Na+(s) + O^{2-}(s)$

$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

ionic equation: $2Na(s) + Cl_2(g) \rightarrow 2Na^+(s) + 2Cl^-(s)$

The modern theoretical definition of oxidation describes chemical reactions that resemble the reaction with oxygen, but with compounds other than oxygen. In both cases, electrons are lost by the sodium atoms.

2. ionic equation: $Sn(s) + Cd^{2+}(aq) + 2NO_3^-(aq) \rightarrow Sn^{2+}(aq) + 2NO_3^-(aq) + Cd(s)$

net ionic equation: $Sn(s) + Cd^{2+}(aq) \rightarrow Sn^{2+}(aq) + Cd(s)$

(a) The spectator ion, $NO_3^-(aq)$, appears unchanged on both the reactant and product side of the equation. Just as a spectator watches a sporting event but does not take part, a spectator ion "watches" a chemical reaction as it does not change during the reaction.

(b) The reducing agent is Sn(s), as it is oxidized in the reaction.

(c) The oxidizing agent is $Cd(NO_3)_2(aq)$, or $Cd^{2+}(aq)$, as it is reduced in the reaction.

(d) Sn(s) is oxidized.

(e) $Cd^{2+}(aq)$ is reduced.

3. The reaction in b) will occur spontaneously.
4. Smelting reduces the iron from iron(II) ions to metallic iron, also called pig iron. The pig iron is then refined in a series of steps that involves the oxidation of the impurities to remove them from the purified steel.
5. **(a)** N = +3

   (b) P = +5

   (c) Si = +4

   (d) N = –3, H = +1, S = +6, O = –2
6. No. An oxidation number of –2 for oxygen in water means that the oxygen atom attracts electrons more strongly than the hydrogen atom in each O—H bond. Oxygen is considered to have "gained" the two electrons it shares with the hydrogen atoms.
7. For reactions that use only molecular compounds, the oxidation number method is easier.
8. **(a)** $Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(\ell)$

   (b) The reaction is not a redox reaction because the oxidation numbers do not change.
9. Sodium chloride melts at temperatures above 800 °C. Therefore, the electrolysis is not carried out under standard conditions.

Answers to Applying Concepts Questions

10. Students should use either the oxidation number method or the half-reaction method (preferable) to write the reaction between acidified hydrogen peroxide, $H_2O_2(aq)$ and aqueous potassium permanganate. The balanced net ionic equation is shown below:

$$5H_2O_2(aq) + 2MnO_4^-(aq) + 6H^+(aq) \rightarrow 5O_2(g) + 2Mn^{2+}(aq) + 8H_2O(\ell)$$

Volume of $MnO_4^-(aq)$ added =

21.54 mL – 1.34 mL = 20.20 mL

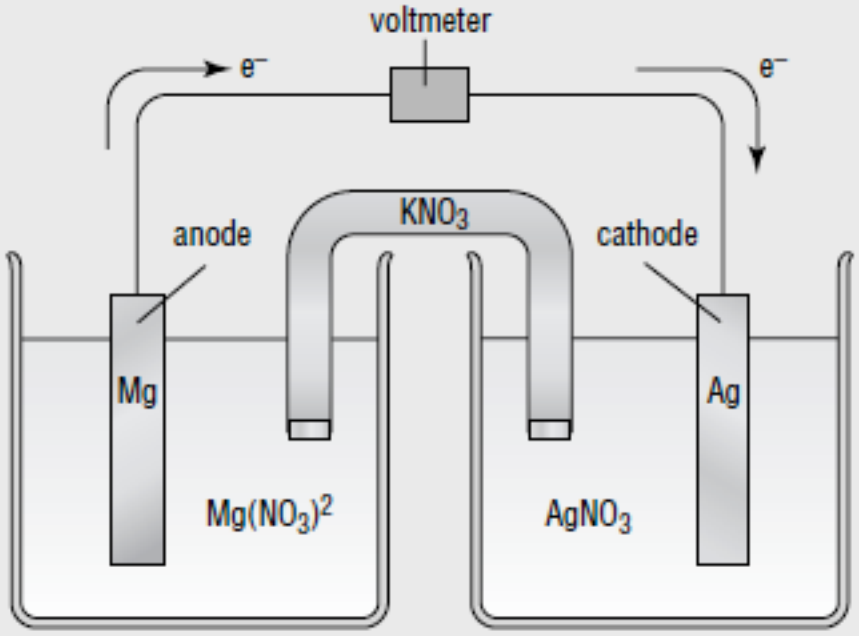
$$C_{H_2O_2} = \frac{0.852\ \text{g}\ KMnO_4(s)}{1} \times \frac{1\ \text{mol}\ KMnO_4(s)}{158.04\ \text{g}\ KMnO_4(s)} \times$$

$$\frac{1}{0.250\ \text{L}} \times \frac{0.0202\ \text{L}}{1} \times \frac{1\ \text{mol}\ MnO_4^-(aq)}{1\ \text{mol}\ KMnO_4(s)} \times$$

$$\frac{5\ \text{mol}\ H_2O_2}{2\ \text{mol}\ MnO_4} \times \frac{1}{0.025\ \text{L}}$$

$$= 4.36 \times 10^{-2}\ \text{mol/L}\ H_2O_2(aq)$$

11. (a) to **(c)**



(d) anode: $Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$

cathode: $Ag^+(aq) + e^- \rightarrow Ag(s)$

(e) $Mg(s) + 2Ag^+(aq) \rightarrow Mg^{2+}(aq) + 2Ag(s)$

(f) $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

$E^\circ_{cell} = 0.80\ \text{V} - (-2.37\ \text{V})$

$E^\circ_{cell} = +3.17\ \text{V}$

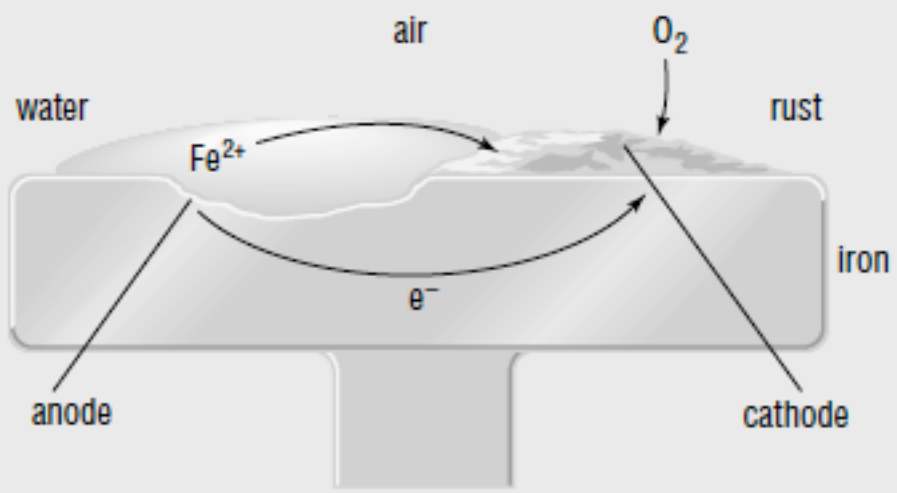
(g) $Mg(s) \mid Mg^{2+}(aq) \parallel Ag^+(aq) \mid Ag(s)$

12. An inert electrode is made from a material that is neither a reactant nor a product of the redox reaction but can carry a current and provide a surface on which redox reactions can occur. Carbon and platinum are commonly used as inert electrodes.

13. "Standard reduction potential" refers to the amount of energy conferred on a unit of charge as the material listed is reduced. Since potential must be measured in reference to another half-cell, the hydrogen half-cell is used as the reference point; its reduction potential is set at 0.00 V and all other half-cells are measured relative to this.

14. A primary battery cannot be recharged while a secondary battery can be recharged.

15.



Anode: $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$

Cathode: $O_2(g) + 2H_2O(\ell) + 4e^- \rightarrow 4OH^-(aq)$

Overall reaction: $2Fe(s) + O_2(g) + 2H_2O(\ell) \rightarrow 2Fe^{2+}(aq) + 4OH^-(aq)$

16. Protect from exposure to oxygen and water: Cover the material with paint, oil, or plastic to prevent exposure to the reactants involved in the corrosion reaction.

Alloys: Mix iron with other metals to reduce corrosion.

Galvanization: Coat with zinc. As the zinc corrodes, a protective layer is formed. Also, zinc is a stronger reducing agent than iron, so it will corrode preferentially.

Cathodic protection: Attach a stronger reducing agent such as magnesium, which will corrode preferentially, protecting the iron.

17. In most cases, electrolysis will lead to the reduction of the strongest oxidizing agent. In an aqueous solution, water will also behave as an oxidizing agent. Water is a stronger oxidizing agent than many reactive metal ions including $Cr^{2+}(aq)$, $Al^{3+}(aq)$, $Mg^{2+}(aq)$, $Na^{+}(aq)$, $Ca^{2+}(aq)$, $Ba^{2+}(aq)$, $K^{+}(aq)$ and $Li^{+}(aq)$, and will therefore be reduced before any of these metal ions.

18. The impure metal is placed at the anode. As a current is applied, the copper on the anode dissolved, and pure copper can be plated on the cathode.

19. The electrolyte is sulfuric acid. As the battery discharges, sulfate ions combine with the lead in the battery to form solid lead(II) sulfate and water. Sulfuric acid solution is denser than water. Therefore, the electrolyte solution is most dense when it is fully charged.

20. The following half-reactions are involved:

(1) $2H^{+}(aq) + 2e^{-} \rightarrow H_2(g)$

(reduction at cathode, $E^{\circ}_{cell} = 0.00$ V)

(2) $2H_2O(\ell) \rightarrow O_2(g) + 4H^{+}(aq) + 4e^{-}$

(oxidation at anode, E°_{cell} for reverse reaction $= 1.23$ V)

(3) $2H_2O(\ell) + 2e^{-} \rightarrow H_2(g) + 2OH^{-}(aq)$

(reduction at cathode, $E^{\circ}_{cell} = -0.83$ V)

(4) $2Cl^{-}(aq) \rightarrow Cl_2(g) + 2e^{-}$

(oxidation at anode, E°_{cell} for reverse reaction $= 1.36$ V)

Using $E^{\circ}_{cell} = E^{\circ}_{cathode} - E^{\circ}_{anode}$ the most positive (least negative) cell potential is obtained using half-reactions (1) and (2):

$E^\circ_{cell} = 0.000\ V - (1.23\ V) = -1.23\ V$

The overall reaction is:

$2H_2O(\ell) \rightarrow O_2(g) + 2H_2(g)$

The predicted products are oxygen and hydrogen. However, the production of hydrogen requires overvoltage. Therefore, chlorine may form instead.

21. **(a)** If the lead completely covers the iron, no oxygen can contact the iron as long as the coating is intact. Lead has a more positive reduction potential than iron, and will oxidize more slowly than iron.

(b) If any iron is exposed, it acts as an anode and is oxidized (rusts).

22. Students could choose to construct an electrolytic cell with an iron nail at the cathode in a solution of any cation listed below it on the reactivity series (such as zinc or aluminium).

23. **(a)** The half-cell that has the most positive reduction potential acts as an anode with the three other half-cells. The half-cell that has the second most positive reduction potential acts as an anode with the two half-cells with more negative reduction potentials. The half-cell that has the third most positive reduction potential acts as an anode with the half-cell with the most negative reduction potential. In total, you could make six different voltaic cells.

(b) You only need to determine three cell potentials. Choose one half-cell as your reference, and measure the cell potential of the reference half-cell paired with each of the other three half-cells. This is analogous to determining standard reduction potentials based on the hydrogen electrode.

24. Refer to Figure 13.23 on page 502 of the student textbook. Replace the zinc electrode with lead and the copper electrode with silver. Use the electrolytes lead nitrate and silver nitrate.

25. Reduction of hydrogen ions:

$2H^+(aq) + 2e^- \rightarrow H_2(g)$ $E^o = 0.00$ V

Reduction of water molecules:

$2H_2O(\ell) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$ $E^o = -0.83$ V

Reverse the first equation and add the two equations:

$2H_2O(\ell) \rightarrow 2H^+(aq) + 2OH^-(aq)$

$E^o_{cell} = E^o_{cathode} - E^o_{anode}$

$= -0.83\text{ V} - (0.00) = -0.83\text{ V}$

The negative sign indicates that the reaction is non-spontaneous under standard conditions. This is consistent with the extremely low concentration of hydrogen and hydroxide ions in pure water.

26. A balanced half-reaction equation indicates whether oxidation or reduction is taking place, how many electrons are involved, and identifies oxidizing and reducing agents and products. For example, consider the following equations:

$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

Because the electrons are on the products side, the half-reaction is an oxidation. Zinc is a reducing agent. Each zinc atom loses two electrons, and Zn^{2+} is produced.

$Cu^{2+} + 2e^- \rightarrow Cu(s)$

Because the electrons are on the reactants side, the half-reaction is a reduction. Copper(II) ions are an oxidizing agent. Each copper ion gains two electrons, and copper metal is produced.

27. (a) Metal A is the more effective reducing agent. It is oxidized and facilitates the reduction of the other metal. For example, in the following reaction, magnesium is the more effective reducing agent.

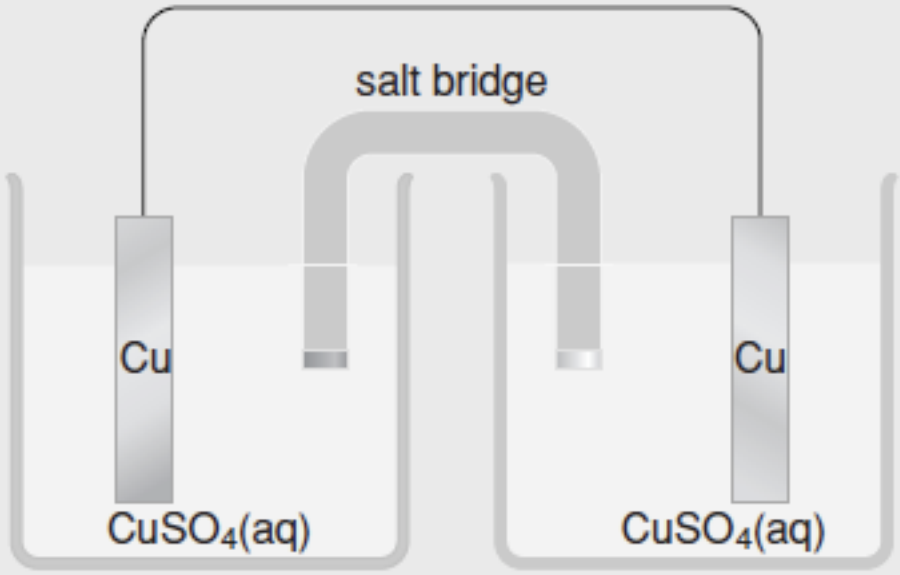
$Mg(s) + Zn(NO_3)_2(aq) \rightarrow Zn(s) + Mg(NO_3)_2(aq)$

(b) For one nonmetal to replace another during a single displacement reaction, it must gain electrons more easily than the nonmetal it is displacing. Therefore, it must be a better oxidizing agent. An example is:

$F_2(g) + 2NaCl(aq) \rightarrow 2NaF(aq) + Cl_2(g)$

28. (a) Since the electrodes are both made of copper, and neither water nor sulfate ions will create a potential difference in the cell, the cell potential is zero.

(b) By adding an external source of electrons to create an anode and a cathode, the cell can be operated as an electrolytic cell. The copper metal anode will decrease in mass and the copper metal cathode will increase in mass.



29. Half-reactions represent oxidations or reductions, which never occur independently. For each electron "lost" in an oxidation reaction, an electron must be "gained" in a reduction reaction. Electrons never appear in the overall equation for a redox reaction, which is consistent with the fact that electrons are not created or destroyed in chemical reactions.

30. (a) A disproportionation reaction occurs when one element is both oxidized and reduced in the same reaction.

(b) $3ClO^-(aq) \rightarrow 2Cl^-(aq) + ClO_3^-(aq)$

31. (a) An electrolytic cell drives a non-spontaneous reaction, meaning that E^o_{cell} is negative. Since $E^o_{cell} = E^o_{cathode} - E^o_{anode}$ and E^o_{anode} is 0 V, E^o cathode must be negative.

(b) Since $E^o_{cell} = E^o_{cathode} - E^o_{anode}$ and $E^o_{cathode}$ is 0 V, E^o_{anode} must be positive for E^o_{cell} to be negative.

32. The Downs cell and the chlor-alkali process are similar in that they both produce chlorine gas through oxidation of chloride ions. In the chlor-alkali process, a saturated aqueous solution of sodium chloride (brine) is electrolyzed. The products are chlorine gas, hydrogen gas, and sodium hydroxide. A Downs cell, in contrast, electrolyzes molten sodium chloride. The products are sodium metal and chlorine gas. To maintain the sodium chloride as a liquid, the reaction in a Downs cell must be carried out at a very high temperature (about 600 °C). This is not necessary in the chlor-alkali process.

33. (a) Students should emphasize that oxidation occurs when an atom "loses" electrons and reduction occurs when an atom "gains" electrons, and that neither process occurs independently of the other. They may include sample half-reactions and redox equations showing oxidation numbers.

(b) Voltaic cells convert chemical energy to electrical energy through spontaneous redox reactions. Electrolytic cells convert electrical energy to chemical energy by providing a source of electrons that allow a non-spontaneous redox reaction to occur. Students should include the idea that voltaic cells are essentially

the reverse of electrolytic cells. Students should provide an example, including a diagram, of each type of cell.

Answers to Solving Problems Questions

34. (a) $Hg_2^{2+}(aq) + 2e^- \rightarrow 2Hg$

(b) $TiO_2(s) \rightarrow Ti^{2+}(aq)$ (acidic conditions)

Titanium is reduced (+4 to +2):

$TiO_2(s) \rightarrow Ti^{2+}(aq)$

Balance the number of oxygen atoms:

$TiO_2(s) \rightarrow Ti^{2+}(aq) + 2H_2O$

Balance the number of hydrogen atoms (acidic conditions):

$4H^+(aq) + TiO_2(s) \rightarrow Ti^{2+}(aq) + 2H_2O$

Balance charge by adding electrons:

$4H^+(aq) + TiO_2(s) + 2e^- \rightarrow Ti^{2+}(aq) + 2H_2O$

(c) $I_2 \rightarrow H_3IO_6^{3-}$ (basic conditions)

Iodine is oxidized (0 to +6):

$I_2(s) \rightarrow H_3IO_6^{3-}(aq)$

Balance the number of iodine atoms:

$I_2(s) \rightarrow 2H_3IO_6^{3-}(aq)$

Balance the number of oxygen atoms:

$12H_2O(\ell) + I_2(s) \rightarrow 2H_3IO_6^{3-}(aq)$

Balance the number of hydrogen atoms:

$12H_2O(\ell) + I_2(s) \rightarrow 2H_3IO_6^{3-} + 18H^+(aq)$

Add $OH^-(aq)$ to neutralize the hydrogen ions:

$$18OH^-(aq) + 12H_2O(\ell) + I_2(s) \rightarrow 2H_3IO_6^{3-}(aq) + 18H^+(aq) + 18OH^-(aq)$$

Combine hydrogen and hydroxide ions and remove water common to both sides. Balance charge by adding electrons:

$$18OH^-(aq) + I_2(s) \rightarrow 2H_3IO_6^{3-}(aq) + 6H_2O(\ell) + 12e^-$$

35. (a) Oxidation

$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$

Reduction

$4H^+(aq) + NO_3^-(aq) + 3e^- \rightarrow NO(g) + 2H_2O(\ell)$

Write the half-reactions using the LCM

$3Fe(s) \rightarrow 3Fe^{2+}(aq) + 6e^-$

$8H^+(aq) + 2NO_3^-(aq) + 6e^- \rightarrow 2NO(g) + H_2O(\ell)$

Add the two half-reactions, replace the spectator nitrate ions, and combine the nitrate ions with the hydrogen ions and the iron(II) ions:

$3Fe(s) + 8HNO_3(aq) \rightarrow 3Fe(NO_3)_2(aq) + 2NO(g) + 4H_2O(\ell)$

(b) Oxidation

$Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^-$

Reduction

$4H^+(aq) + NO_3^-(aq) + 3e^- \rightarrow NO(g) + H_2O(\ell)$

Write the half-reactions using the LCM

$3Fe^{2+}(aq) \rightarrow 3Fe^{3+}(aq) + 3e^-$

$4H^+(aq) + NO_3^-(aq) + 3e^- \rightarrow NO(g) + 2H_2O(\ell)$

Add the two half-reactions, replace the spectator nitrate ions, and combine the nitrate ions with the hydrogen ions and the iron(III) ions:

$3Fe(NO_3)_2(aq) + 4HNO_3(aq) \rightarrow 3Fe(NO_3)_3(aq) + NO(g) + 2H_2O(\ell)$

(c) Oxidation

$Fe(s) \rightarrow Fe^{3+}(aq) + 3e^-$

Reduction

$4H^+(aq) + NO_3^-(aq) + 3e^- \rightarrow NO(g) + 2H_2O(\ell)$

The balanced reaction is:

$Fe(s) + 4H^+(aq) + NO_3^-(aq) \rightarrow Fe^{3+}(aq) + NO(g) + 2H_2O(\ell)$

(d) These all represent corrosion reactions.

36. (a) The anode is an inert carbon electrode, and the cathode is a silver electrode. The positive electrode is the cathode, the negative electrode is the anode.

(b) oxidation at anode: $2I^-(aq) \rightarrow I_2(s) + 2e^-$

reduction at cathode: $Ag^+(aq) + 1e^- \rightarrow Ag(s)$

overall cell reaction: $2I^-(aq) + 2Ag^+(aq) \rightarrow I_2(s) + 2Ag(s)$

(c) Silver is the oxidizing agent and the iodide ion is the reducing agent.

(d) $E^o_{cell} = E^o_{cathode} - E^o_{anode}$

$= 0.80\ V - (+0.54\ V) = +0.26\ V$

37. (a) The reaction is not a redox reaction because the oxidation numbers do not change.

$Cl_2O_7(aq) + H_2O(\ell) \rightarrow 2HClO_4(aq)$

(b) The oxidizing agent is $ClO_3^-(aq)$ and the reducing agent is $I_2(s)$. Students should use the half-reaction method to balance the equation for acidic conditions.

$3I_2(s) + 3H_2O(\ell) + 5ClO_3^-(aq) \rightarrow 6IO_3^-(aq) + 5Cl^-(aq) + 6H^+(aq)$

(c) Br_2 is the oxidizing agent and S^{2-} is the reducing agent. Students should use the half-reaction method to balance the equation for basic conditions.

$8OH^-(aq) + S^{2-}(aq) + 4Br_2(\ell) \rightarrow SO_4^{2-}(aq) + 8Br_2(\ell) + 4H_2O(\ell)$

(d) The oxidizing agent is HNO_3 and the reducing agent is H_2S. Students can balance this equation using the half-reactions method or the oxidation number method.

$3H_2S(aq) + 2HNO_3(aq) \rightarrow 3S(s) + 2NO(g) + 4H_2O(\ell)$

38. (a) $E^{o}_{cell} = E^{o}_{cathode} - E^{o}_{anode} = -0.23$ V; non–spontaneous

(b) $E^{o}_{cell} = E^{o}_{cathode} - E^{o}_{anode} = +0.70$ V; spontaneous

(c) $E^{o}_{cell} = E^{o}_{cathode} - E^{o}_{anode} = +0.42$ V; spontaneous

39. (a) Students should balance the equation for acidic conditions using the half-reaction method:

$$2MnO_4^-(aq) + 5C_2O_4^{2-}(aq) + 16H^+(aq) \rightarrow 2Mn^{2+}(aq) + 10CO_2(g) + 8H_2O(\ell)$$

(b) Amount of $KMnO_4$ = 14.28 ml × 0.1575 mol/1000 mL

$= 2.249 \times 10^{-3}$ mol

Amount $(NH_4)_2C_2O_4$ = 5 mol $Na_2C_2O_4$/2 $KMnO_4$ × 2.249×10^{-3} mol $KMnO_4$

$= 5.623 \times 10^{-3}$ mol $Na_2C_2O_4$

Concentration of oxalate solution = 5.623×10^{-3} mol $Na_2C_2O_4$/0.02500L

= 0.225 mol/L

40. Cerium(IV) ions will be reduced and tin(II) ions will be oxidized according to the following equations:

Reduction: $Ce^{4+}(aq) + e^- \rightarrow Ce^{3+}(aq)$

Oxidation: $Sn^{2+}(aq) \rightarrow Sn^{4+}(aq) + 2e^-$

Based on these equations, 1 mol of cerium(IV) ions are reduced by 0.5 mol of tin(II) ions.

Amount of cerium(IV) ions = 0.024 38 L × 0.2113 mol/L

$= 5.151 \times 10^{-3}$ mol cerium(IV) ions

Amount of Sn(II) ions = 2.576×10^{-3} mol

Mass of Sn = 2.576×10^{-3} mol × 118.71 g/mol

= 0.305 8 g Sn

Mass % of tin = 20.73%

41. (a) $$\text{Amount of magnesium ions} = \frac{3.38\ \text{g}}{24.31\ \text{g/mol}}$$
$$= 0.139\ \text{mol}$$
$$\text{Amount of electrons} = 0.139\ \text{mol}\ Mg^{2+} \times \frac{2\ \text{mol}\ e^-}{1\ \text{mol}\ Mg^{2+}}$$
$$= 0.278\ \text{mol}\ e^-$$
$$\text{Quantity of electricity} = \frac{96\,500\ \text{C}}{1\ \text{mol}\ e^-} \times 0.278\ \text{mol}\ e^-$$
$$= 2.68 \times 10^4\ \text{C}$$

(b) $$2.68 \times 10^4\ \text{C} = 100 \times 10^{-3}\ \text{A} \times \text{t}$$
$$\text{t} = 2.68 \times 10^5\ \text{s} \div 3600\ \text{s/h}$$
$$= 74.5\ \text{h}$$

42. $$\text{Quantity of electricity} = 2.92\ \text{A} \times 71.0\ \text{min} \times \frac{60\ \text{s}}{1\ \text{min}}$$
$$= 1.24 \times 10^4\ \text{C}$$
$$\text{Amount of electrons} = 1.244 \times 10^4\ \text{C} \times \frac{1\ \text{mol}\ e^-}{96\,500\ \text{C}}$$
$$= 0.1289\ \text{mol}\ e^-$$

$$\text{Amount of aluminium} = 0.1289\ \text{mol}\ e^- \times \frac{1\ \text{mol Al}}{3\ \text{mol}\ e^-}$$
$$= 0.042\,97\ \text{mol Al}$$
$$\text{Mass of aluminium} = 0.0429\ \text{mol Al} \times \frac{26.98\ \text{g}}{\text{mol Al}}$$
$$= 1.16\ \text{g}$$

43. $$\text{Quantity of electricity} = 5000\ \text{A} \times 7\ \text{days} \times \frac{24\ \text{h}}{1\ \text{day}} \times \frac{3600\ \text{s}}{1\ \text{h}}$$
$$= 3.024 \times 10^9\ \text{C}$$
$$\text{Amount of electrons} = 3.024 \times 10^9\ \text{C} \times \frac{1\ \text{mol}\ e^-}{96\,500\ \text{C}}$$
$$= 3.134 \times 10^4\ \text{mol}\ e^-$$

$$\text{Amount of fluorine} = 3.134 \times 10^4 \text{ mol } e^- \times \frac{1 \text{ mol } F_2}{2 \text{ mol } e^-}$$
$$= 1.567 \times 10^4 \text{ mol } F_2$$

$$\text{Mass of fluorine} = 1.567 \times 10^4 \text{ mol } F_2 \times \frac{38.00 \text{ g}}{\text{mol } F_2}$$
$$= 5.954 \times 10^5 \text{ g or } 0.5954 \text{ t}$$

44. (a) (i) The gas at the anode would be chlorine gas, $Cl_2(g)$, as long as sufficient voltage is supplied (this is the chlor-alkali cell).

anode: $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$

(ii) A "pop" test could be used to identify the gas in tube C. A lit splint would ignite the hydrogen gas, $H_2(g)$, produced, resulting in a pop.

cathode: $2H_2O(\ell) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$

(iii) Platinum can also be used as an inert electrode.

(iv) $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

$E^\circ_{cell} = 0.83 \text{ V} - (-1.36 \text{ V})$

$E^\circ_{cell} = 2.19 \text{ V}$

(v) Collector A contains chlorine gas, which is made at the same rate as hydrogen gas, therefore equal volumes will be made. Chlorine is very soluble in water, however, so you will not see as much gas collected in the tube.

(vi) Litmus will turn blue at the cathode since hydroxide ions are produced. At the anode, the litmus will be bleached by the presence of chlorine gas dissolved in the water.

(b) (i) Chlorine gas will still be produced at the anode, but at the cathode, solid sodium will be deposited:

anode: $2Cl^-(\ell) \rightarrow Cl_2(g) + 2e^-$

cathode: $Na^+(\ell) + e^- \rightarrow Na(s)$

(ii) $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

$E^\circ_{cell} = 2.71\ V - (-1.36\ V)$

$E^\circ_{cell} = 4.07\ V$

The calculated minimum voltage uses "standard" reduction potentials, which are recorded based on 1.0 mol/L aqueous solutions at 0 °C and 101.325 kPa. In order to operate this cell, the temperature is raised to 600 °C, and calcium chloride is added to the molten salt. This calculation assumes that the change in temperature and the use of molten salts instead of aqueous solutions will not affect the cell potential.

(iii) Cell 1: $2H_2O(\ell) + 2Cl^-(aq) \rightarrow Cl_2(g) + H_2(g) + 2OH^-(aq)$

Cell 2: $2Cl^-(\ell) + 2Na^+(\ell) \rightarrow Cl_2(g) + 2Na(s)$

In cell 1, water is reduced while in cell 2, sodium ions are reduced.

(c) (i) When water is present, it will be preferentially reduced compared to sodium ions because it is a much stronger oxidizing agent than sodium ions. To reduce sodium ions, stronger oxidizing agents cannot be present.

(ii) Aqueous salts such as sodium nitrate, $NaNO_3(aq)$, or sodium sulfate, $Na_2SO_4(aq)$, would lead to the disproportionation of water and lead to oxygen gas production at the anode and hydrogen gas production at the cathode.

(iii) The flow of electrons is always from the anode to the cathode.

Answers to Making Connections Questions

45. Since the speed of developing is related to the rate of decomposition of the silver halide, and since silver iodide is used for high-speed film, silver iodide is more sensitive to light than silver bromide.

46. (a) $8Al(s) + 3NH_4ClO_4(aq) \rightarrow 4Al_2O_3(s) + 3NH_4Cl(aq)$

The oxidizing agent is ammonium perchlorate. The reducing agent is aluminium.

(b) $2H_2(g) + O_2(g) \rightarrow 2H_2O(\ell)$

(c) In the first reaction, ammonium chloride is produced, which could participate in harmful reactions in the environment. The second reaction produces only water. Therefore, the second reaction is more environmentally friendly.

47. Corroding iron mains must be repaired or replaced, which costs money. Corroding mains may also leak and their hydraulic capacity may decrease, which decreases their efficiency. Leaking and corroding water mains may allow contaminants into drinking water. Several ways to combat these problems include replacing iron water mains with polymer or cement mains, which do not corrode, or using sacrificial anodes to prevent corrosion.

48. Iron is abundant (the second most abundant metal after aluminium), relatively inexpensive to mine and process, and can be used to create a number of different alloys with a variety of useful physical properties.